Chapter 11: Chemical Bonding

I. Introduction to Chemical Bonding

a. Types of chemical bonding

i. A **chemical bond** is a mutual attraction between nuclei and the valence electrons of different atoms that binds the atoms together

ii. You could, for shorthand, just say that a bond is the force that holds the atoms to one another

iii. Why do atoms bond? That's like asking why a coffee mug will hit the ground if you let go of it: the cup will be in a state of lower potential energy (will be more stable) if it hits the ground than if it doesn't (is left floating in the air). If, when two atoms bond, they are higher in energy than the original separate atoms, then such a molecule is unlikely to form.

iv. **Ionic bond:** when metals and nonmetals bond.

1. Metals tend to lose electrons
2. Nonmetals tend to gain electrons
3. When a negative particle (electron) is lost, the remaining atom is positively charged, and is therefore an ion. **Positive ion: "cation"**
4. When an atom gains an electron it becomes a negative ion (an **anion**)
5. When an anion and a cation are near each other, they will become attracted due to their opposite charges. This is an ionic bond.

v. **Covalent bond**: when atoms of nonmetals bond.

1. Because both atoms have high electronegativities (want to take the electrons), the atoms end up sharing the electrons.

vi. **Metallic bonds**: bonds between metal atoms. Because all of the atoms in the bond have low electronegativities, the electrons are "pooled" between the atoms.

b. Electronegativity

i. An ionic bond can also be defined as when the electronegativity difference between two atoms is 1.6 or greater.

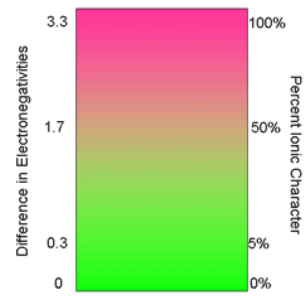
Group / Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
6	Cs 0.79	Ba 0.89	Lu 1.27	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo

Periodic table of electronegativity using the Pauling scale

ii. Electronegativity is the tendency of an atom to attract electrons to itself when it is bonded to another atom. Noble gases have zero or near-zero electronegativities. Metals have low electronegativities, and nonmetals other than noble gases have high electronegativities.

iii. The difference in electronegativity between two atoms can be used to classify the type of bond that they form

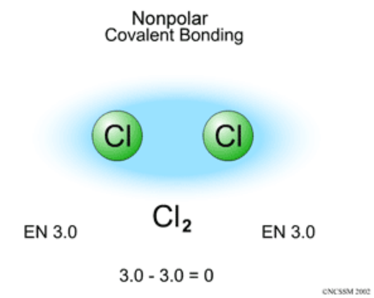
1. **Nonpolar covalent**: 0 to 0.4
2. **Polar covalent**: 0.5 to 1.7
3. **Ionic**: 1.8 and greater



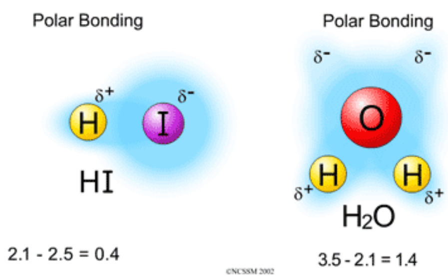
c. Ionic bonding

d. Covalent bonding

i. nonpolar covalent bonding



ii. polar covalent bonding



II. Covalent Bonding and Molecular Compounds

a. **<u>The Octet Rule</u>**: Atoms lose gain, or share electrons so as to have a completely filled highest-occupied energy level.

i. For our purposes, this means that the atom will try to have a total of 8 highest energy level s and p electrons

ii. These electrons are called the **<u>valence electrons</u>**. They are the electrons that are involved in bonding. These electrons are also called the outer shell electrons.

iii. (Contrast the valence electrons with the core electrons, which are lower in energy and not involved in bonding.)

iv. The number of valence electrons varies regularly across the main group elements:

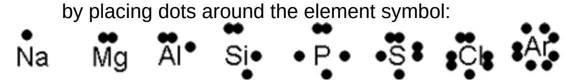
Group	1	2	(T.M., I.T.M.)	13	14	15	16	17	18
Valence e^-'s	1	2		3	4	5	6	7	8 (but not He!)

b. Exceptions to the octet rule

i. Helium and elements close to it also satisfy the octet rule by achieving a full outer shell. However, in the case of the first shell (a.k.a. the 1st energy level) 2 electrons constitute a full shell.

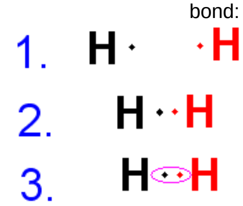
ii. Atoms beyond the 2nd period can use unoccupied d sublevels to achieve "expanded" octets. Examples: SF_6, PCl_5. More on this later.

c. Electron dot notation: The number of valence electrons is indicated by placing dots around the element symbol:



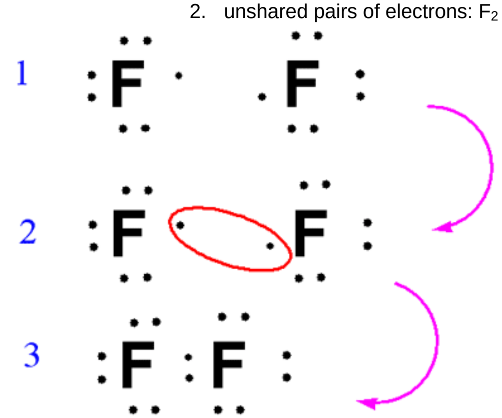
d. Lewis structures

i. Let's use dot structures to show the sharing of electrons that occurs between H and H when they form a covalent bond:



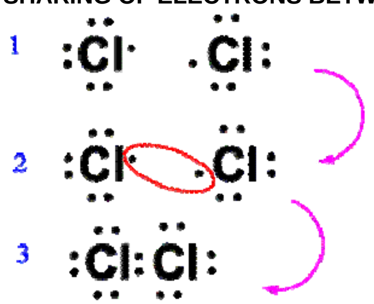
ii. Let's do the same thing for F and F

1. shared pairs of electrons: F_2 has 1 shared pair
2. unshared pairs of electrons: F_2 has 6 shared pairs



iii. Cl and Cl (That's the element **chlorine**, not Carbon and Iodine!)

SHARING OF ELECTRONS BETWEEN Cl AND Cl



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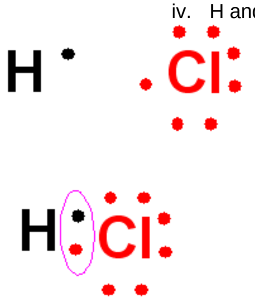
iv. H and Cl



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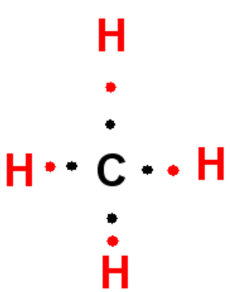
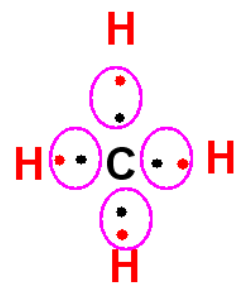
v. C and four H's make methane:



Then . . . . join the electrons in order to make 4 single covalent bonds.



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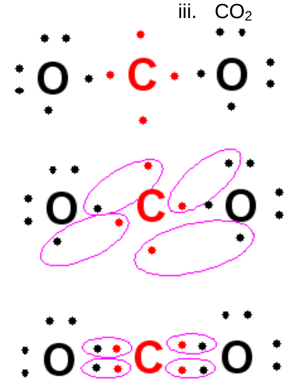
e. Multiple covalent bonds

i. Sometimes two or more pairs of electrons need to be shared between electrons in order for the octet rule to be satisfied

ii. O_2 (To keep it simple, we will only consider the case that obeys the octet rule – the diamagnetic case.)

iii. CO_2



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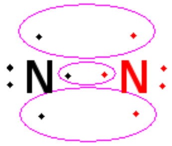
iv. C_2H_4

v. N_2



vi. C_2H_2

vii. Single Covalent Bond: one pair of electrons shared between atoms

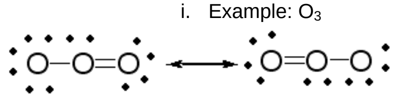
viii. Double covalent bond: two pairs of electrons shared between two atoms. Not the same as two single bonds. Not twice as strong as two single bonds

ix. Triple covalent bond: three pairs of electrons shared between two atoms. Not the same as three single bonds. Not three times as strong as three single bonds

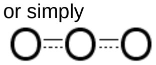
x. Bond lengths and bond strengths

f. Resonance Structures: when one or more possible dot structures for a molecule is possible, it is helpful to think of all of these structures as contributing to the final structure.

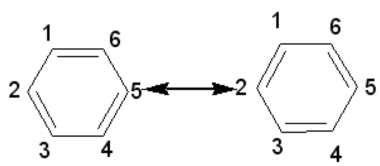
i. Example: O_3



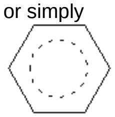
or simply



ii. Example benzene: C_6H_6

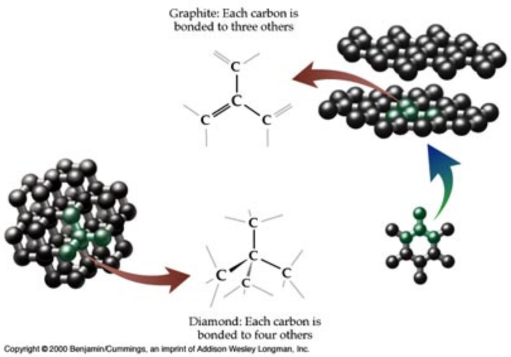


or simply



g. Covalent network bonding

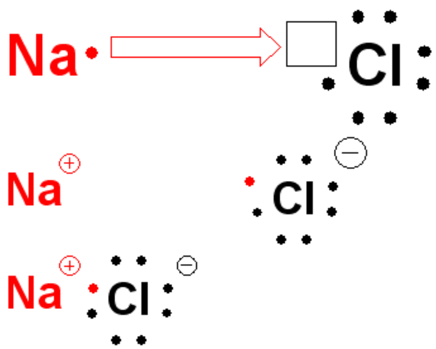
i. Example: diamond – this is a model of diamond's network solid structure. Diamond is one **allotrope** of the element carbon that occurs in nature. (When two or more forms of an element occur in nature, each of these forms is called an allotrope of that element.) Another common allotrope of carbon is graphite, which is still pure carbon but the atoms are linked to one another differently than they are in diamond.



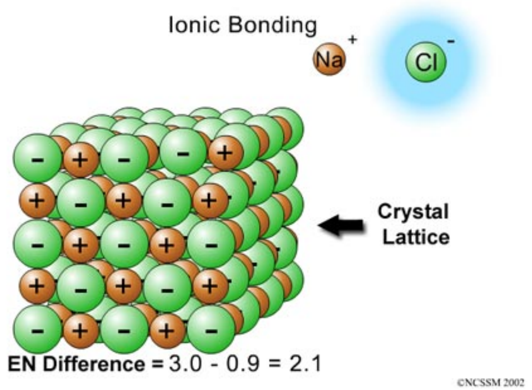
III. Ionic Bonding and Ionic Compounds

a. Formation of ionic compounds

TRANSFER OF ELECTRONS from Na to Cl



b. Characteristics of ionic bonding



c. Comparing ionic and molecular compounds

IV. Metallic Bonding

a. The "sea of electrons model"

b. Summary: comparing ionic, covalent, and metallic bonding

V. Molecular Geometry

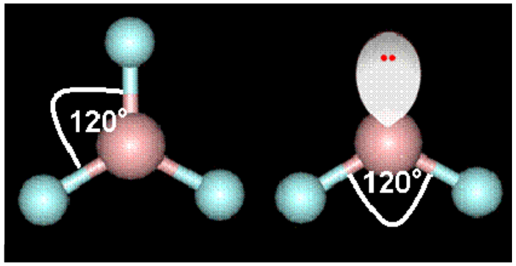
a. VSEPR theory

i. <u>**V**</u>alence <u>**S**</u>hell <u>**E**</u>lectron <u>**P**</u>air <u>**R**</u>epulsion <u>**T**</u>heory

ii. All electrons in the valence shell (highest energy level) travel in pairs

iii. These pairs (shared pairs as well as lone pairs) will repel each other and determine the molecule's geometry

Electron domain geometry: Trigonal Planar:



Molecular Geometry: **Trigonal Planar** **Bent**

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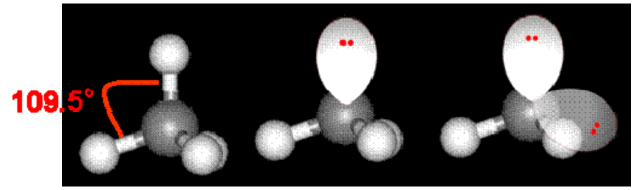
Electron domain geometry: Tetrahedral:



Tetrahedral **Trigonal pyramidal:** **Bent**

Molecular Models Lab

Molecule	Dot Structure	Structural Formula	Shape	Polar Bonds?	Polar Molecule?
H_2					
Br_2					
H_2O					
CO_2					
PF_3					
SO_2					
SO_4^{2-}					
O_2					
O_3					
N_2					
CH_4					
CO_3^{2-}					
C_2H_6					
C_2H_4					
C_2H_2					
* BF_3					
* PCl_5					
* SF_6					
* NO_2					
C_6H_{14}					
** C_6H_{12}					
** C_6H_6					

b. Intermolecular forces of attraction

i. **Hydrogen bonding**

1. Molecules that have an electronegative atom with one or more unshared pairs of electrons and also have an electropositive hydrogen can undergo hydrogen bonding

2. H-bonding is a temporary covalent bond between an unshared pair on one atom and an electropositive H-atom on <u>another</u> molecule.

3. An H-bond is about 5% as strong as a normal covalent bond.

ii. Van der Waals forces

1. Dipole-dipole interactions

a. Polar molecules (also known as dipoles) will align themselves positive end-to-negative end. This makes it more difficult to remove one molecule from another. For instance, the boiling point of a polar substance will be higher than the boiling point of a nonpolar substance.

2. London dispersion forces

a. What about nonpolar molecules? If octane (gasoline – C_8H_{18}) and other hydrocarbons aren't polar and can't undergo hydrogen bonding, then why aren't all of these substances gases? For instance, gasoline is a liquid. Wax is a solid.

b. The answer is that these molecules form <u>temporary</u> dipoles which attract one another.

c. This is similar to dipole-dipole interactions, but the force of attraction is much smaller